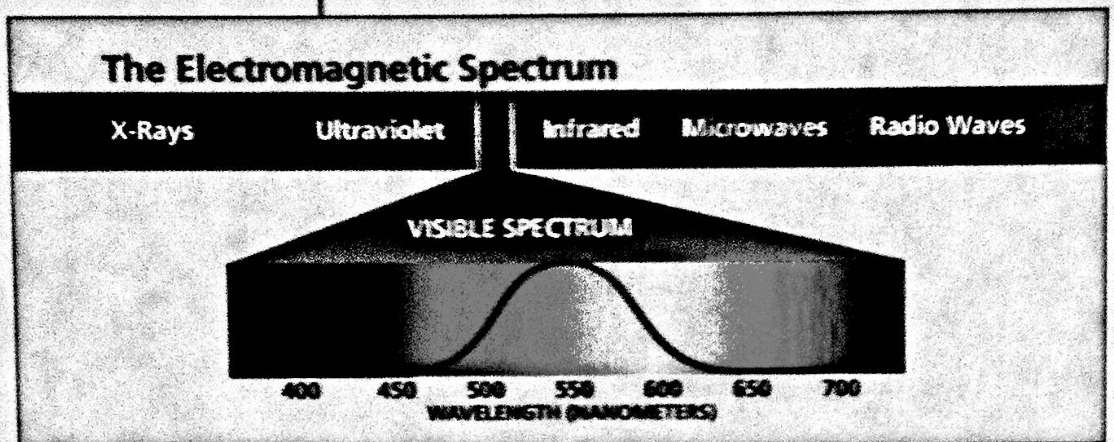
BOHR'S MODEL OF AN ATOM

The Electromagnetic Spectrum

Visible light is only one part of the **electromagnetic spectrum.** You've probably heard of some of the other parts including ultraviolet, infrared, x-rays, gamma rays, microwaves, and radio waves. As you demonstrated in calculations using Bohr's model, the light from some of the transitions is in the ultraviolet. Infrared light is also emitted as the electron jumps from E_4 to E_3 and E_5 to E_3 and other higher energy levels.



Many people do not think of radio waves as being light waves, but they are part of the electromagnetic spectrum. The problem is that you can only see a very small part of the electromagnetic spectrum. Often, you hear radio announcers say that they are broadcasting at a certain **frequency.** Your FM radio dial may have MHz (megahertz) printed on the side. This tells you that the numbers correspond to frequencies in units of MHz or 10^6 Hz. Frequency tells you the number of cycles or waves that are being produced per second. The unit for frequency is a hertz (Hz). 1 Hz = 1 cycle/s = 1 s^{-1}. Normally, frequency is read as per second and the cycles are dropped from the terminology.

Wavelength (λ), where λ is the Greek letter *lambda*, is the distance from crest to crest of a wave. All light waves travel at the same speed. The speed of electromagnetic radiation is constant and it is called the speed of light (c). $c = 2.998 \times 10^8$ m/s or 3.00×10^8 m/s. From this information you can calculate the frequency of light of a given wavelength. The equation that is used for this is:

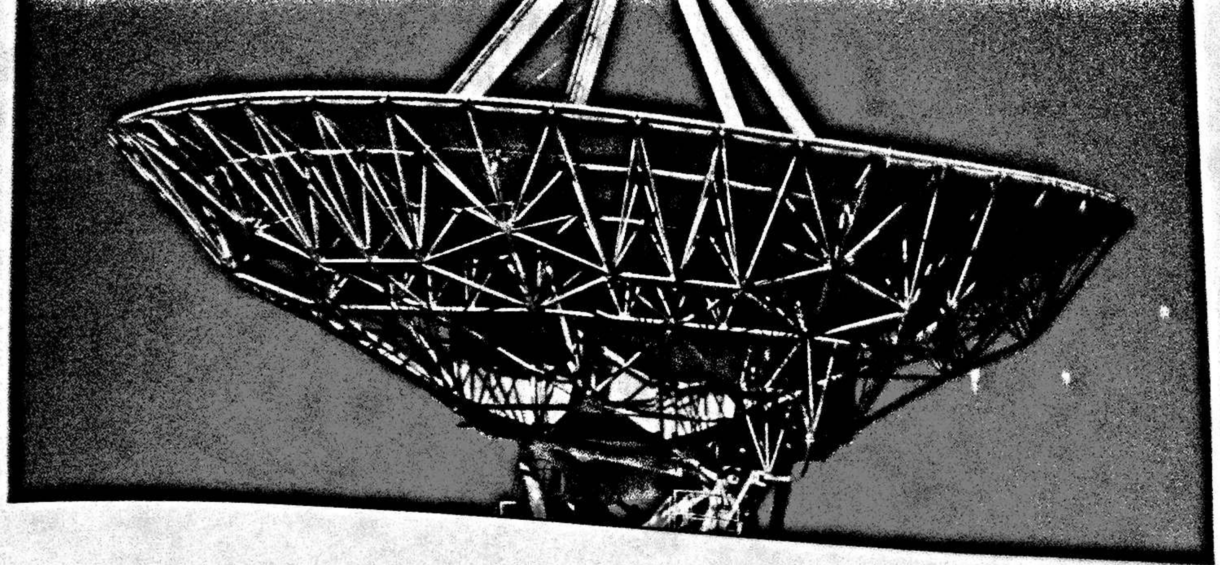
$$f = \frac{c}{\lambda}$$

Chem Words

electromagnetic spectrum: the complete spectrum of electromagnetic radiation, such as radio waves, microwaves, infrared, visible, ultraviolet, x-rays, and gamma rays.

frequency: the number of waves per second or cycles per second or hertz (Hz).

wavelength: the distance measured from crest to crest of one complete wave or cycle.



As an example, if the wavelength is 434.2 nm, then the frequency is:

$$f = \frac{2.998 \times 10^{8}\ \text{m/s}}{434.2 \times 10^{-9}\ \text{m}}$$

$$= 6.905 \times 10^{14}\ \text{cycles/s or } 6.905 \times 10^{14}\ \text{Hz.}$$

As you go across the electromagnetic spectrum you should note that the wavelength continues to get smaller as the frequency increases. Also, you should understand that the energy of the spectrum increases as you go from radio waves to x-rays or gamma rays. Max Planck, a German physicist, found that the energy of a wavelength could be calculated. The equation that he developed was based on measuring the change in energy from one level to another level like you did in the **Investigate** section. The equation he developed is

$E = hf$,

where h is Planck's constant and is $6.63 \times 10^{-34}\ \text{J} \cdot \text{s}$
and f is the frequency.

The corresponding energy of the red light above would be:

$$E = hf$$

$$= (6.63 \times 10^{-34}\ \text{J} \cdot \text{s})\ (4.567 \times 10^{14}\ \text{Hz})$$

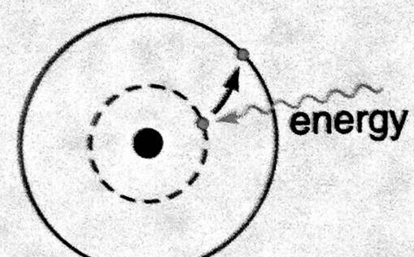
$$= 3.03 \times 10^{-19}\ \text{J.}$$

So the next time that you are standing around a campfire, you can inform your fellow campers that red light has less energy than blue light and you can also tell them how to calculate these values. ☺

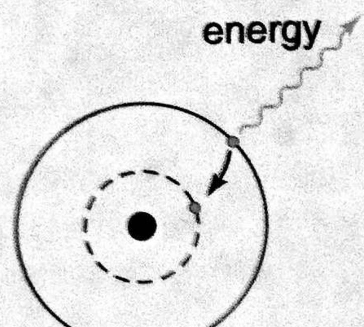
Bohr's Atomic Theory

Niels Bohr, a brilliant Danish physicist, was aware that his theory of electron jumps had incredible success but also raised some problems.

Bohr proposed a "planetary" model of the atom. He theorized that electrons travel in nearly circular paths, called **orbits**, around the nucleus. Each electron orbit has a definite amount of energy, and the farther away the electron is from the nucleus, the greater is its energy. Bohr suggested the revolutionary idea that electrons "jump" between energy levels (orbits) in a quantum fashion. That is, they can never exist in an in-between state. Thus, when an atom absorbs or gives off energy (as in light or heat), the electron jumps to higher or lower orbits. Electrons are the most stable when they are at lower energy levels closer to the nucleus. Bohr's theory could only account for the spectrum of hydrogen and not for the spectra of any other element. Bohr's theory could not explain why only certain orbits were allowed, nor how the electron could jump from one orbit to another. Other scientists improved on Bohr's model as they discovered more about the atom and quantum mechanics.



(a) An electron gains a quantum of energy.



(b) An electron loses a quantum of energy.

Chem Words

orbit: the path of the electron in its motion around the nucleus of Bohr's hydrogen atom.

Checking Up

1. How are visible light, ultraviolet light, infrared light, x-rays, gamma rays, microwaves, and radio waves related?
2. Explain the meaning of wavelength.
3. Why is "planetary" model an appropriate name for Bohr's model of the atom?
4. How do the energy levels of different electron orbits compare?
5. Why do elements produce certain color light when heated?